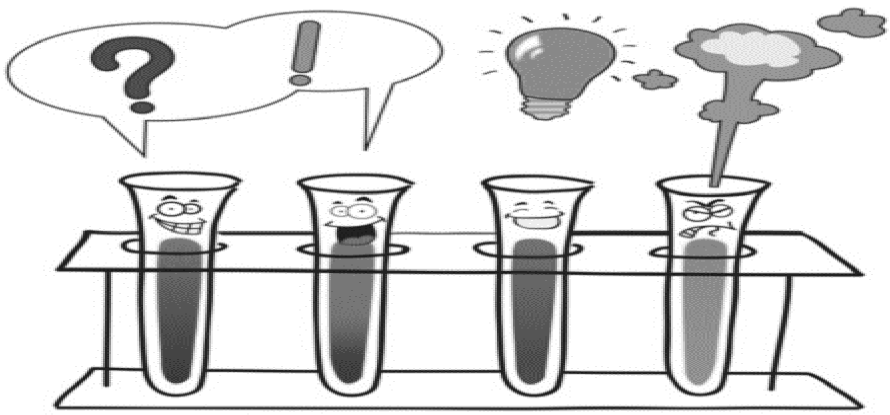
Chemistry

Second Secondary

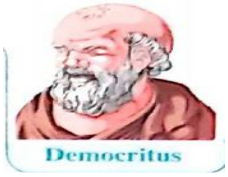
2024 / 2025



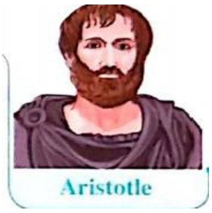
Chapter one

Structure of the atom

Atomic structure

1- Greek philosophers: (Democritus) 

Any piece of matter can be divided into smaller parts and each part can be subdivided into smaller parts which can't be divided this part is called Atom.

2- Aristotle: 

a-He refused the idea of Greek philosophers about the atom .

b- He supposed that all matters composed of 4 constituents which are (water, air , dust and fire and postulated that the cheap metals such as iron or copper can be changed into precious ones like gold by changing the percentage of four constituents.

3- Boyl (lrish 1661): 

i-He refused the Aristotle concept.

ii-He was the 1st scientist to define the element as pure simple substance which can not be analvsed into smple one by normal chemical methods.

4) Dulton s atom (Englsh 1803) :

He supposed thet :

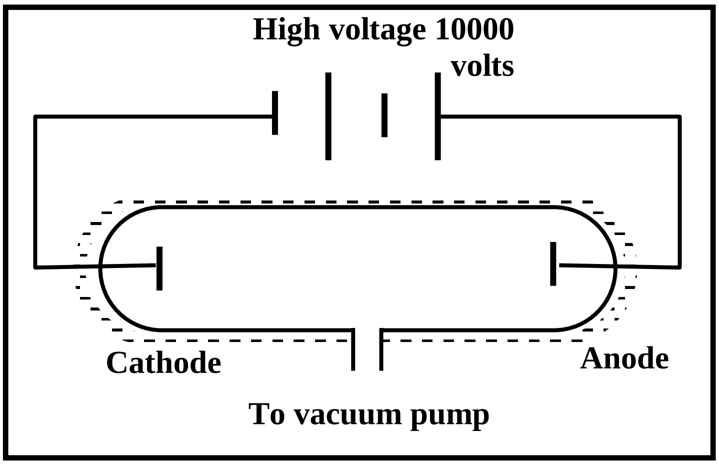
1- **The substance consists of very amall particles called atoms.**
2- **Every element consists of very small dense atoms which can t be divided.**
3- **Atoms f the same element are identical.**
4- **Atome of different are different.**

N.B **By the 19 th century**

Scientists had accepted the idea that elements consists of atoms but they knew very little about those atoms.

Cathode- ravs experiment

(discovery of the electron)



a- All gases under normal conditions of pressure and temp (76 cm. hg. 25c) don t conduct electricity .

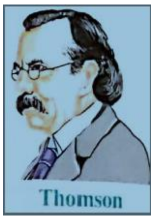
b- If a glass tube evacuated from the gas to decrease its pressure to reach 0.01.1 ------------- 0.001 m.mhg

The gas will conduct electric current.

c- If the potential difference between the tow poles increases up 10.000 Volts a Flow of invisble rays are emitted from the cathode causing glowing to the wall of tube behind the anode and called cathode ray .

Properties of cuthode rays

1) Consists of tiny particles have mass and velocity.
2) Transfers in straigh lines glowing the glass facing the cathode.
3) Have negatve charge.
4) Have a thermal effect.
5) Affected by electric and Magnetic field.
6) Cathode rays don t change by changing either cathode material or type of the gas which proves that cathode rays take part in the structure of all substances.

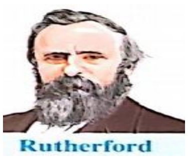


5 – Thomson s atom 1897 :

He conclude from the last experiment that

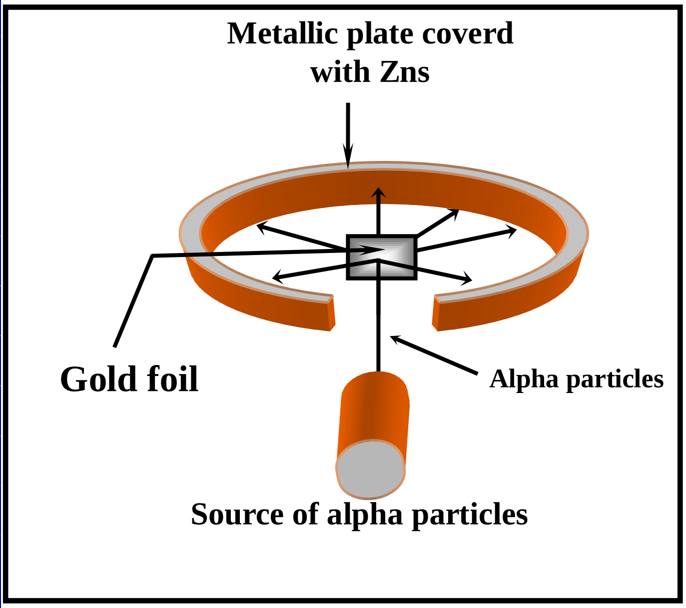
i- The atom is a homogenous sphere of positive electricity.

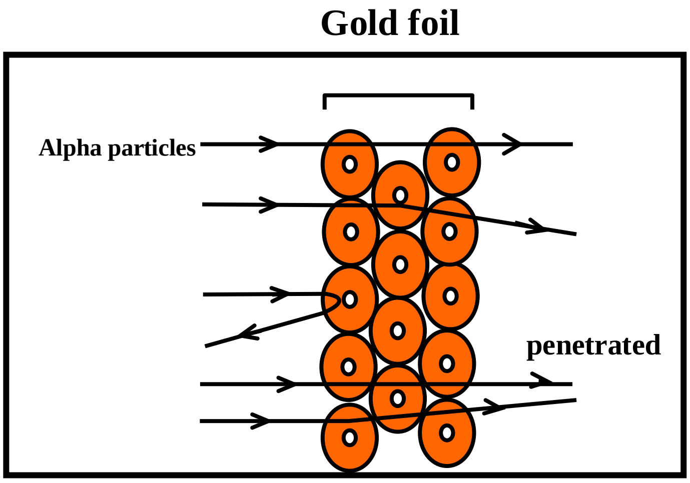
ii- Inside it there are negative electrons enough to make it electrically neutral.



Rutherford s experiment

In 1911 Geiger and Marsden performed a famous exp According to suggestion of Ruther ford by the followng apparatus.





1 – He allowed alpha particles to hit a mtallic plate lined with Zinc suphide (glows when hits with alpha rays)

2 – On placing a gold foil in the front of alpha rays he concluded the following from the following observation.

Observation	Result
1- Most of alpha particles appeared in the same position before putting gold foil.	1- Most of the atom is a space not solid as explained by Dalton and thomson.
2- A very small percentage of alpha particles reflected back to appears as flashes on other side of sheet.	2- The atom has very small part with very small volume but high density.
3- Some flashes appeared on the sides of 1^{st} site.	3- The dense part of the atom which concentrate in it most mass have same charge of alpha particle (+ve) which called nucleus of the atom.

Structure of the atom

Atomic No:

No of protons inside the nucleus or no of electrons inside the E. levels.

Mass No:

Sum of protons and neutrons inside the nucleus.

Isotopes:

Are different forms of atoms of the same element which have same atomic nos but different mass no.

Ex: hydrogen has 3 isotopes: ……….

Rutherford:

Rutherford was the 1st scientist who stated the concept of the atomic structure.

1-The Atom:

Although it has very small size but it has a complicated structure that resembles the solar system in which electrons revolve around the central nucleus in orbits as planets revolve around the sun.

2-The Nucleus:

Is much smaller than the atom. Located in the centre of the atom with (+ve) charge. There is a big space between the nucleus and orbits of electrons, so most of the atom is a space. Most mass of the atom is concentrated in the nucleus as mass of e is very small and can be neglected.

3-Electrons:

1-Have negligible mass compared to that of the nucleus.
2-No of electrons (-ve) equals no of ptotons (+ve) so the atom is electrically neutral.
3-Electrons revolve around the nucleus in a fixed orbit as electrons are affected by two forces equal in strength but in opposite direction, which are :

a- Force of attraction of the nucleus to electrons.

b- Centrifugal force due to velocity of electron around the nucleus.

Give reason: Electrons are not attracted to the nucleus.

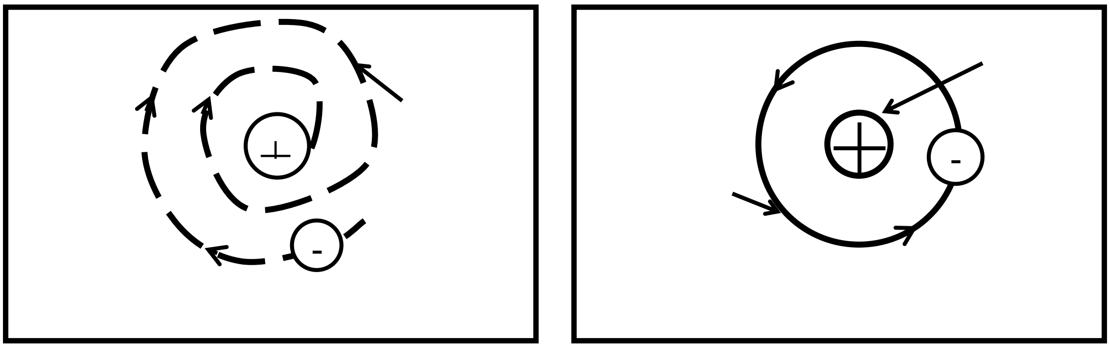
Explain: Structure of the atom in the view of rutherford.

Objections on Rutherford's atomic model (Maxwell's theory):

Rutherford's concept was contradicted by maxwell's theory (Which was based on laws of newton mechanics and concerned with the movement of relatively large bodies).

- Which states that: " When an electrecally charged body moves in orbit, it will lose its energy gradually by emission of radiation causing gradual decrease in orbit radius".

By applying this theory on electron movement in Rutherford's atom, we would expect that electrons are in a state of continuous emission of radiation, so the atomic radius will decrease and electrons move in a spiral orbit until they hit the nucleus.



Give reason: Contradiction between the classical mechanical laws and rutherford.

Bohr's Atomic Model

Bohr's postulates:

1- A positively charged nucleus exists in the center of the atom.
2- Atom is electrically neutral as no of p^+ s equals to no of e's.
3- Electrons revolve around the nucleus in orbits due to centrifugal and attraction forces.
4- Electrons orbit the nucleus in a rapid movement without gaining or losing energy.
5- Electrons orbit the nucleus only in a definite allowed energy levels, so they can't be found at intermediate distance.

6- Each electron in the atom has a definite amount of energy depending on the distance between is E.L and nucleus. This energy increases as its radius increases.
7- It was found that the maximum no of energy levels in the heaviest known atoms in their ground state (unexcited) is only seven (K, L, M, N, O, P, Q). Each level has energy expressed by a whole no called principle Q. No.

Ex: The 1st E. Level K its principle Q. no = 1
The 2nd E. level L its principle Q. no = 2

8- If when atom is excited by heating (Quantum) or by electric discharge the electron will transfer to a higher E. level agrees with the absorbed quantum. The excited electron in the higher E. level is then unstable, so it returns to its original level losing the same quantum of energy, which it gained during excitation in the form of radiation have definite wavelength and frequency.

❖ Remarks:

1- The quantum: Is the amount of energy gained or lost when an electron jumps from one E. level to another.
2- The difference in energy between levels (Q) is not equal i.e. the difference in this energy decreases further from the nucleus. This means that the quantum of energy required to transfer an electron from one energy level to another is not equal.
3- The electron does not move from its level to another unless the energy absorbed or emitted is equal to the difference in energy between 2 levels i.e. one quantum.
(There is no half quantum for instance). Q can't be divided or doubled

Give reason: It is wrong to say that e' to be transferred from E.L (K) to E.L (M) needs amount of energy equals 2 quantum.

Excited Atom:

- It is an atom that acquired an amount of energy (Q) sufficient to transfer its e's from their original E.L to higher ones.

Advantages of Bohr:

Bohr's atomic theory succeeded in the following ways:

1- It explained hydrogen atom spectrum.
2- He introduced the idea of quantum no to detect energy of electrons in energy levels.
3- He proved that electrons during that electrons during rotation around the nucleus in ground state do not radiate energy, so they will not fall back to the nucleus .

(a reconcilion between Rutherford and Maxwell).

Disadvantages of Bohr's theory:

1 – If failed to explain the spectrum of any other element even that of He except hydrogen (Simplest Electronic System).
2 – He considered the electron as a (-ve) particle only and did not consider that it also has wave properties.
3 – He postulated that it is possible to determine precisely both speed and location of an electron at the same time. This is experimentally impossible because the apparatus used will change either speed or location of electron so the result won't be accurate.
4 – He described the electron when moving in a circular planer orbit, which means that hydrogen atom is planer. Later it was confirmed that hydrogen atom has 3 dimensional co-ordinates.

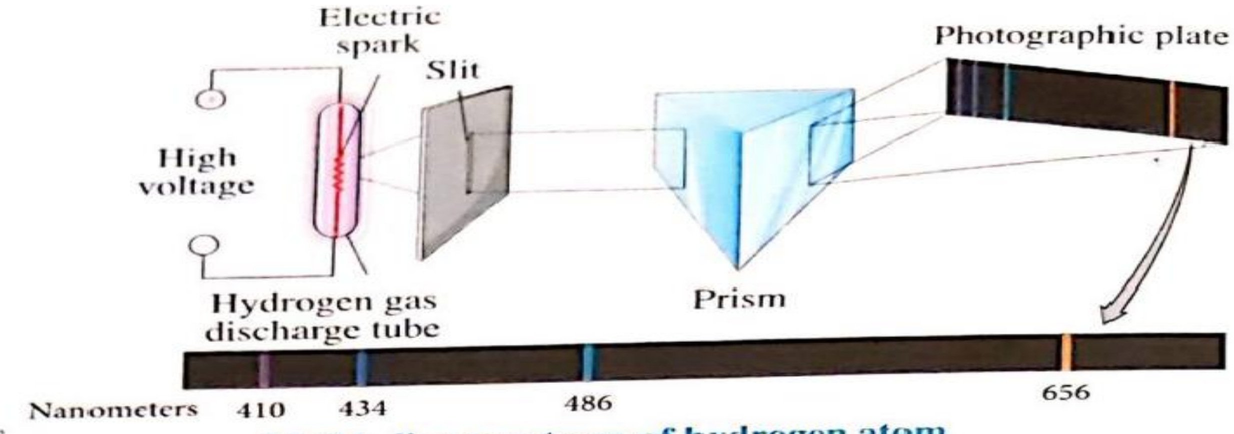
Bohr's Theory

The atomic spectrum

Studying and explaining the atomic spectrum was the key to he atomic structure in 1913 and deserved noble prize at 1922.

Atomic emission spectrum:

1 – By heating gases or vapours of substances to a high temperature (by heat or electricity) under low pressure it produces light.
2 – By using spectroscope we find that this light consists of a fixed number of coloured lines called line spectrum.



Visible line spectrum of hydrogen atom

Line spectrum of any element:

By exp. Proved that spectrum line differs from one element to another like finger prints.

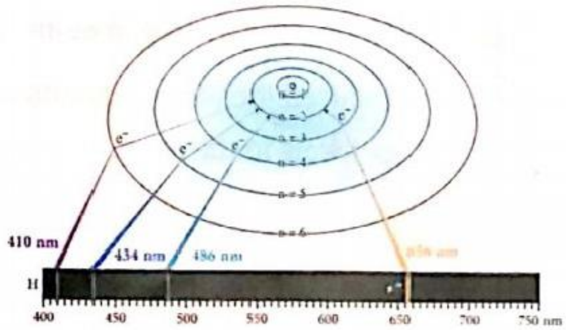
N.B:

Line spectrum of sun rays shows that composed of <u>hydrogen and Helium.</u>

For illustration only

The opposite figure which is representing the spectral line of the hydrogen atom doesn't represent the electron transferring from :

- The different energy levels to the first energy level, because the wavelength of the emitted ray from the excited electron is located in the invisible region of the ultraviolet rays.
- The seventh energy level to the second energy level, because the wavelength of the emitted ray from the excited electron is located in the invisible region of the infrared rays.



The visible spectral line of hydrogen atom consists of four colored lines.

The principles of Modern Atomic Theory:

1 – Dual nature of electron.

2 – The Heisenberg uncertainty principle.

3 – Finding the mathematical expression which describes the wave motion of electron, its shape and its energy.

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1 – The dual nature of the electron

The experimental data showed that the electron has a dual nature i.e

a) It is a material particle. b) It has wave properties.

* De Broglie principle:

Every moving body (such as electron or the nucleus of an atom or whole molecule)is associated with (accompanied by) a wave motion (or matter waves) which has some properties of light waves.

The matter wave motion differs from electromagnetic waves n ………..

Matter Waves	Electromagnetic Wave
1- They are not separating from the moving body	1- They are separated from the moving body
2- Their speed is not equal to the speed of light	Their speed is equal to the speed of light

2- The Heisenberg uncertainty principle: (quantum mechanics)

It is practically impossible to determine both position and the velocity of the electron exactly (precisely) at the same time. We can only say that it is probably to a greater r lesser extent to locate the electron in this or in that place. This is to speak in terms of probability seems to be more precise.

3. The wave equation for motion of electron inside the atom:

(Schrodinger wave mechanics theory): He applied the ideas of Blanck, Einestein, De Broglie and Heisenberg so he shows that:

- ❖ It is possible to determine the allowed energy levels of the electron and define the region of space around the nucleus where it is most probable to find the electron in each energy level.
- ❖ The electronic motion around nucleus has a wave properties therefore the position to use the term electron cloud to describe any orbital.

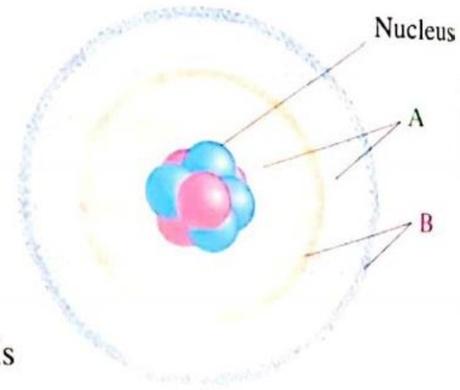
Electron Cloud: (used to describe any orbital)

"Area of space around the nucleus where there is a great probability for finding electrons in all direction and all positions."

Electron cloud

It is the region of space around the nucleus, in which the electron probable exists in all directions and distances (dimensions).

- There are regions inside the electron cloud in which probability of finding the electron increases, each of them is termed by "the orbital" (region B).



Electron cloud and orbital

Orbital

It is the region within the electron cloud of high probability of finding the electron.

❖ **The difference between the orbit and orbital concepts according to both Bohr and the wave mechanics theories:**

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<tr><th>Bohr s theory</th><th>Schrodinger s wave mechanics theory</th></tr>
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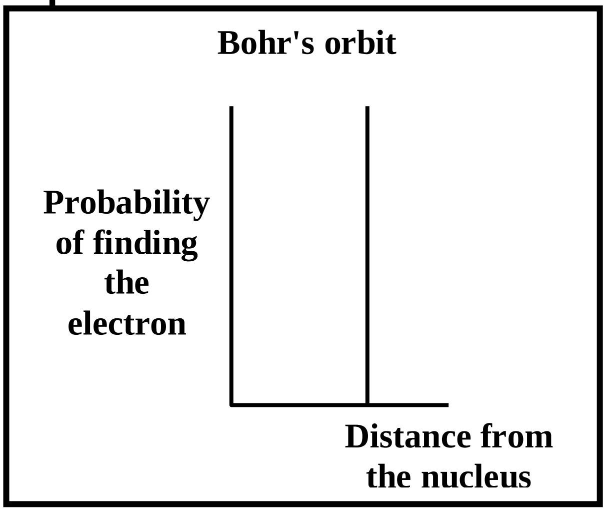
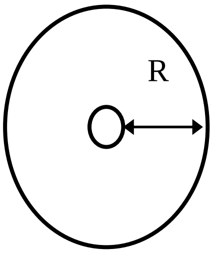
- It is a circular planer orbit with particular radii





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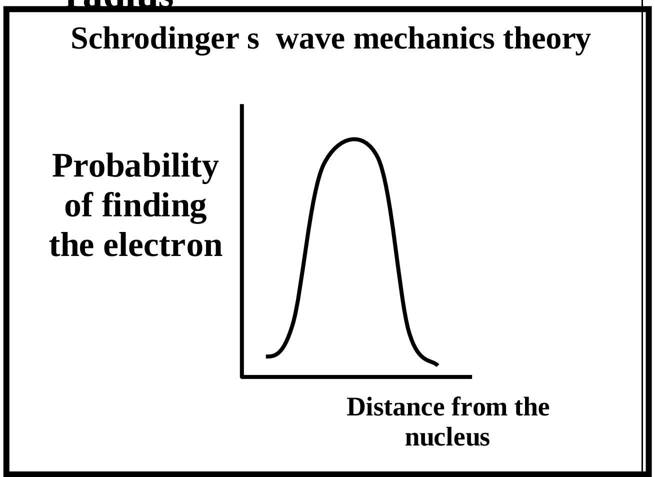
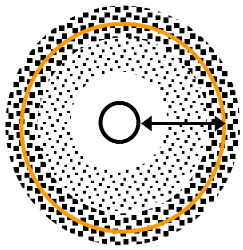
- It is an electron cloud used to describe any orbital
- Electron cloud the regions of high density of dots represent the region of high probability of finding the (e) from which it is possible to define the atomic radius





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❖ **The mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers.**

- Quantum Numbers:

These nos define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they define the energy, shape and direction of orbitals.

1- Principle Q.no (n).
2- Subsidiary or orbital (azimuthal) Q. no (L).
3- Magnetic Q. no (m).
4- Spin Q. no (ms).

1- Principle Q No (n):

Bohr had used this no to define the following:

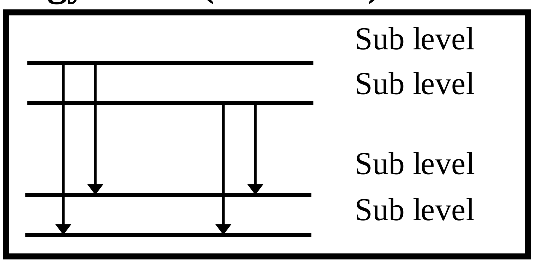
1- Order of principle E. levels their number in the heaviest known atom in the ground state is seven.
2- No of electrons required to fill a given E. level = two times the square of the level no (2n2).

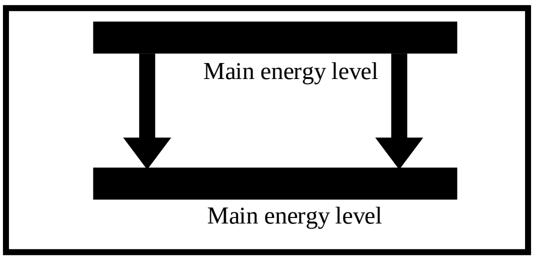
-1st E.L	K	Is filled with	2 electrons
-2nd E.L	L	Is filled with	8 electrons
-3rd E.L	M	Is filled with	18 electrons
-4th E.L	N	Is filled with	32 electrons

- But this rule does not apply to the last three levels (O, P, Q). However, the atom becomes unstable if no of electrons exceeds 32 electrons on any level.

2. Subsidiary Q No (L):

1- Used to detect the no of sub levels in each E. level.
2- The energy sub levels take the symbols s, p, d, f. this is shown by the scientist Somerfield. When he used a spectroscope which has a high resolving power, he found that the single line (which represents electron transition between two different energy levels) is indeed a number of fine spectral lines which represents electron transition between very near energy levels (sublevels).





3- No of sublevels in each energy level = order of principle energy level (n).

-1st E.L	K	has 1 sub level	1s
-2nd E.L	L	has 2 sub level	2s, 2p
-3rd E.L	M	has 3 sub level	3s, 3p, 3d
-4th E.L	N	has 4 sub level	4s, 4p, 4d, 4f

N.B:

- **Energy of sub levels of same E. level is not equal.**

f>d>p>S

- **Energy of same sublevels but of different E. levels also differ in energy.**

Ex: 4d>3d 4p>3p>2p

- **There is a small difference in the energy between sub-levels.**

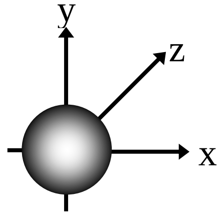
3. Magnetic Q No (m):

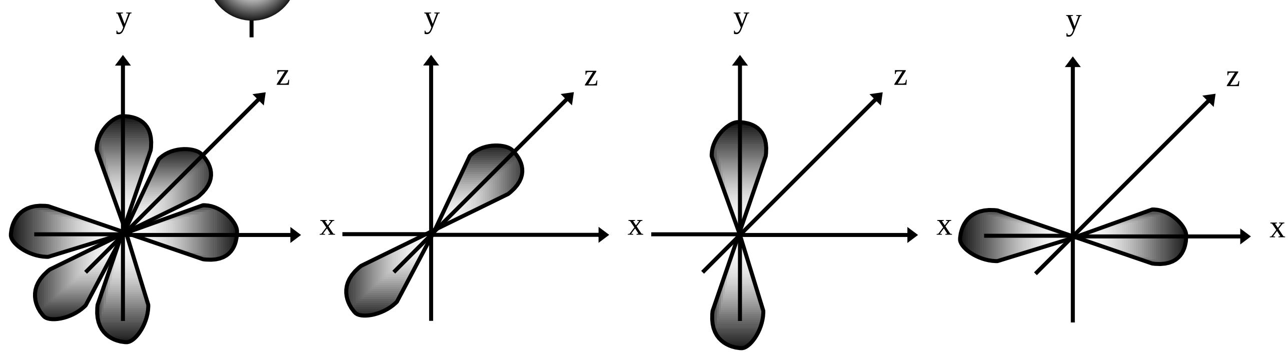
Detected by Zieman when he exposed spectral line to strong magnetic field, he found that each line divides into many lines, so he concluded that each E. sublevel has no orbitals.

- ***Magnetic Q No is characterized by:***

1- Used to detect no of orbitals in each E. sub level and their direction in space.
2- Sublevel (S) has one orbital of spherical symmetrical shape.
3- Sublevel(P) has 3 orbitals.

- **Each orbital (P_x,P_y,P_z) is perpendicular to the other two.**
- **Also the electron cloud of each orbital takes the form of 2 pears meeting head to head (dumb – bell shaped) at a node i.e point of zero electron density.**





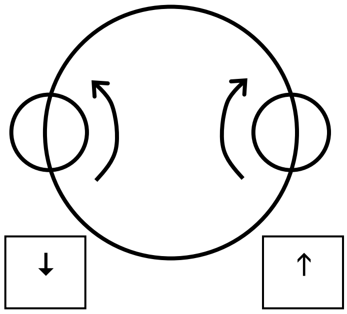
4- Sublevel (d) has 5 orbitals.
5- Sublevel (f) has 7 orbitals.
6- Orbitals of the same sub level are equal in energy and shape.

Ex: $p_x = p_y = p_z$

7- No of orbitals in each E. level = square order of E. level = n^2.

4- <u>Spin Q. No (ms):</u>

- **Detects the direction in which the electron spins around its axis during its rotation around the nucleus.**
- **Each orbital can be saturated by 2 electrons, one electron spins around its axis clockwise while the other electron spins anti – clockwise in order to from 2 opposite magnetic fields to decrease the force of repulsion between them which keep the atom stable.**
- **It has only two possible values + 1/2 – 1/2**



<u>Give reason</u>: Each orbital carries 2 electrons although they are negatively charged.

<u>Summary of the relationship between the principle E.L, sub levels orbitals and no of electrons:</u>

1. **No of energy sublevels = order of principle level (n).**
2. **No of orbitals within a principle level = square the no of the level (n^2).**
3. **No of electrons occupying a given E. level = two times the square order of this level ($2n^2$).**

Ex:

Order of E.L.	E.L	No of sub-levels (n)	No of orbitals (n2)	No of electrons (2n2)
1st	K	1 (1s)	1	2
2nd	L	2 (2s, 2p)	4	8
3rd	M	3 (3s, 3p, 3d)	9	18
4th	N	4 (4s, 4p, 4d, 4f)	16	32

❖ **Principle of distributing electrons:**

- **Aufbau (building – up) principle:**

"Electrons occupy energy sublevels in an ascending order according to increasing energy where the lowest energy sublevel is filled 1st".

Ex: 4s is filled before 3d as energy of 4s < 3d.

K	1S			
L	2S			2p
M	3S			3p
N	4S		3d	4p
O	5S		4d	5p
P	6S	4F	5d	6p
Q	7S	5F	6d	

Examples:

Na 11	$1s^2$	$2s^2$	$2p^6$	$3s^1$				
Ca 20	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$		
Zn 30	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	

- Hund's rule:

State that: "No electron pairing takes place in a given sublevel until each orbital contains one electron."

Ex:

$_7N$	$1s^2 - 2s^2 - 2p^3$

$_8O$	$1s^2 - 2s^2 - 2p^4$

		p_x	p_y	p_y
2p		↑	↑	↑
2s	⥮			
1s	⥮			

		p_x	p_y	p_y
2p		⥮	↑	↑
2s	⥮			
1s	⥮			

- **Atom is stable when the outer sub-level is half completely filled with e's.**

❖Remarks:-

1- **Electrons are preferred to be unpaired before pairing because according to Hund's rule on pairing electrons in the same orbitals, they will repel decreasing stability of the atom.**
2- **Electrons prefer to be paired with another electron than to transfer to a higher sub-level, as the energy needed to transfer it to a higher sub-level.**
3- **Also the spin of single electrons must be in the same direction because this gives the atom more stability.**

Another E. configurations (to the nearest noble gas)

Ex:

$_{3}Li$: (He_2) ,$2S^2$

$_{11}Na$: (Ne_{10}) , $3S^1$

$_{19}K$: (Ar_{18}), 4S1

Questions

I- Choose:

1-

Each of the following is among Dalton's theory postulates, except that ....

(a) atoms of the elements contain protons, neutrons and electrons.
(b) the masses of the atoms of the same element are similar.
(c) the atom is indivisible.
(d) each element is formed of tiny particles called atoms.

2-1st scientist defines the element is ……………

a) Dalton b) Rutherford c) Boyl d) Thomson

1- Substance composed of 4 components which are (water, air, dust and fire) was the idea of ……….

a) Bohr b) Rutherford c) Dalton d) Aristole

2- To prove that cathode rays taking part in all substances ……………

a) Have thermal effect.
b) Transfer in straight line.
c) Have tiny porticles.
d) Don't changing substance of cathode or kind of gas.

3- On heating gases or vapours under low pressure at high temperature

a) Absorbs light. b) Gives light.
b) Gives gamma rays. D) Gives alpha rays.

4-

The ratio of the number of hydrogen atoms to that of nitrogen atoms in ammonia molecule is 3:1, this is consistent with one of the postulates of .... theory.

(a) Thomson's (b) Rutherford's
(c) Bohr's (d) Dalton's

5-

What is the mass ratio of carbon [C = 12] to hydrogen [H = 1] in methane CH_4 ? ....

(a) 1 : 4 (b) 3 : 2
(c) 3 : 1 (d) 4 : 1

6-

Electrical neutrality was first mentioned in ....

(a) Democritus's concept of matter.
(b) Dalton's atom.
(c) Boyle's concept of matter.
(d) Thomson's atom.

7-

Rutherford's model of atom ....

(a) is the recently accepted model of atom.
(b) assumed that the atom is solid.
(c) explained the unique atomic spectrum of the different elements.
(d) assumed that the charge of the electrons equals the charge of the nucleus.

8-

The gold foil experiment which is carried out in Rutherford's lab ....

(a) confirmed Thomson's atomic theory.
(b) is the base for Dalton's theory.
(c) led to discovering the nucleus of the atom.
(d) used in it a source of beta particles.

9-

In Rutherford's experiment, the ratio of the number of alpha particles which are deviated to that of alpha particles which bounce back is ....

(a) more than 1
(b) less than 1
(c) equal 1
(d) infinite number.

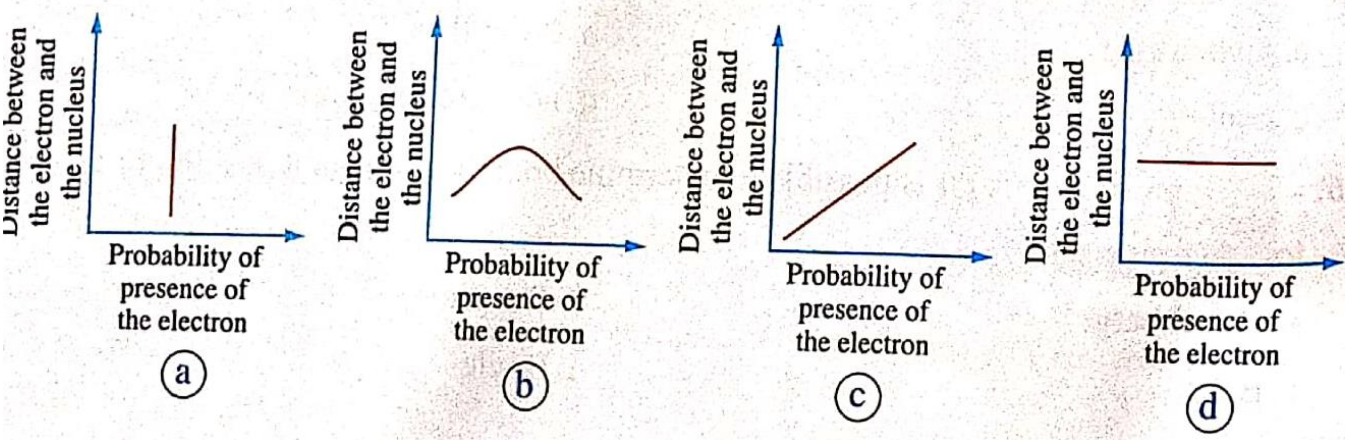
10-

According to Bohr's theory, the orbit in which the electron revolves can be determined through ....

(a) the electron mass.
(b) the electron energy.
(c) the electron charge.
(d) the nucleus charge.

11-

Which of the following graphical figures represents Bohr's concept of the orbit ? ....



12-

On comparing the position of the electron in its ground state, with its position in the excited state, it is ....

a) in the second energy level.

b) in the nucleus.

c) closer to the nucleus.

d) farther from the nucleus.

13-

Which of the following transfers of the electron in hydrogen atom results in the emission of visible light ? ....

a) $(n = 5) \longrightarrow (n = 2)$.

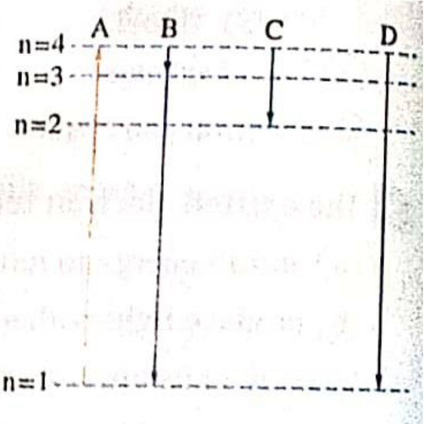
b) $(n = 3) \longrightarrow (n = 1)$.

c) $(n = 5) \longrightarrow (n = 3)$.

d) $(n = 6) \longrightarrow (n = 3)$

14-

The opposite figure represents four energy levels in an excited hydrogen atom, which of the following choices indicates a spectral line in hydrogen atom ? ....



a) A

b) B

c) C

d) D

15-
The difference in energy between each two consecutive energy levels ....

(a) increases by moving away from the nucleus.

(b) has no definite relation.

(c) is perfectly equal.

(d) decreases by moving away from the nucleus.

16-
Bohr's atomic model ....

(a) suggested that the electron occupies a definite energy level only.

(b) explained the line spectrum of hydrogen atom only.

(c) predicted the different energy levels in different multi-electron atoms.

(d) (a) and (b) together.

17-
Bohr's atomic model differs from that of Rutherford, this difference is obvious in Bohr's postulate that the electron ....

(a) produces a spectral line when it loses a quantum.

(b) is a negatively charged particle.

(c) does not produce a spectral line when it loses a quantum.

(d) revolves around the nucleus in certain orbits.

18-
Each of the following is among the properties of the electron, except that it ....

(a) is a material particle.

(b) has wave properties.

(c) loses energy when it transfers from one energy level to a higher level.

(d) deflects by the effect of a magnetic field.

19-
In which of the following sublevels the two quantum numbers of the last electron are ($n = 2$, $\ell = 0$) ? ....

(a) *2s* (b) *2p* (c) *1s* (d) *3p*

20-

The electron which has the two quantum numbers, ($n = 3$, $m_l = +2$) must have the value ....

(a) $m_s = +\frac{1}{2}$ (b) $l = 1$ (c) $l = 0$ (d) $l = 2$

21-

Which of the following quantum numbers values represent an electron in one of the orbitals of $3p_x$ sublevel ? ....

(a) $n = 3$, $l = 2$, $m_l = -1$
(b) $n = 3$, $l = 0$, $m_l = 0$
(c) $n = 3$, $l = 0$, $m_l = +1$
(d) $n = 3$, $l = 1$, $m_l = -1$

22-

The two orbitals ($2s$, $2p_x$) can be similar in ....

(a) the energy.
(b) the shape.
(c) the number of electrons in each of them.
(d) the spatial orientation.

23-

The electrons of $5d$ sublevel in one of the atoms cannot have the magnetic quantum number ....

(a) +1 (b) −1 (c) +2 (d) +3

24-

Number of orbitals of the sublevel which has the values ($n = 3$), ($l = 2$) is ....

(a) 2 (b) 3 (c) 5 (d) 7

25-

Which of the following quantum numbers include a mistake ? ....

(a) $n = 2$, $l = 1$, $m_l = +1$
(b) $n = 4$, $l = 2$, $m_l = +1$
(c) $n = 3$, $l = 3$, $m_l = -2$
(d) $n = 3$, $l = 0$, $m_l = 0$

26-

Which of the following quantum numbers include a mistake ? ....

(a) $n = 6$, $l = 3$, $m_l = +2$
(b) $n = 3$, $l = 2$, $m_l = 0$
(c) $n = 4$, $l = 0$, $m_l = -3$
(d) $n = 3$, $l = 1$, $m_l = -1$

27-

The values of the spin quantum number in the principal energy level in the atom of any element become different when the number of electrons is .... the number of orbitals.

(a) double (b) half (c) equal (d) quarter

28-

Which is easier, losing an electron from $3d$ or from $4s$? ....

(a) $4s$ is more easy as it is closer to the nucleus than $3d$

(b) $4s$ is less easy as it is closer to the nucleus than $3d$

(c) $4s$ is more easy as it is farther from the nucleus than $3d$

(d) $4s$ is less easy as it is farther from the nucleus than $3d$

29-

Which of the following sets of energy sublevels is ordered ascendingly according to the energy ? ....

(a) $4d > 5p = 4f$ (b) $3p = 4s < 3d$

(c) $4p > 4s = 3d$ (d) $5p = 4f > 3d$

30-

Which of the following represent the possible quantum numbers of the last electron of vanadium atom $_{23}V$? ....

(a) $n = 3$, $l = 2$, $m_l = 0$, $m_s = +\frac{1}{2}$

(b) $n = 3$, $l = 2$, $m_l = 0$, $m_s = -\frac{1}{2}$

(c) $n = 4$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$

(d) $n = 4$, $l = 0$, $m_l = +1$, $m_s = -\frac{1}{2}$

31-

The 19th electron in the atom of chromium $_{24}Cr$, has the quantum numbers ....

(a) $n = 3$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$ (b) $n = 3$, $l = 2$, $m_l = -2$, $m_s = +\frac{1}{2}$

(c) $n = 4$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$ (d) $n = 4$, $l = 1$, $m_l = -1$, $m_s = +\frac{1}{2}$

32-

What is the number of orbitals which are completely filled with electrons in an element atom whose atomic number is 16 ? ....

ⓐ 1 ⓑ 7 ⓒ 8 ⓓ 9

33-

What is the atomic number of the element whose electrons occupy 8 orbitals ? ....

ⓐ 8 ⓑ 14 ⓒ 15 ⓓ 26

34-

What is the number of electrons of the last principal energy level of the element which contains 15 completely filled and 2 half filled orbitals ? ....

ⓐ 2 ⓑ 3 ⓒ 4 ⓓ 5

35-

The last energy sublevel in an element contains 3 orbitals which are X, Y and Z, only X contains one electron, and the sum of its $(n + \ell)$ equals 5, what is the atomic number of this element ? ....

ⓐ 19 ⓑ 31 ⓒ 33 ⓓ 41

36-

What is the number of electrons of the penultimate energy level of the element whose atomic number is 28 ? ....

ⓐ 2 ⓑ 8 ⓒ 14 ⓓ 16

37-

What is the electronic configuration of the outermost (third) energy level of a stable atom has 7 valence electrons ? ....

ⓐ $3s^1, 3p^6$ ⓑ $3s^1, 3p^4, 3d^2$ ⓒ $3s^2, 3p^5$ ⓓ $3s^2, 2p^4, 3d^1$

38-

The last sublevel in X^{3+} ion is $2p^6$, what is the number of the half filled orbitals in the atom X ? ....

ⓐ Zero ⓑ 1 ⓒ 2 ⓓ 3

39-

What is the electronic configuration which represents an excited atom ? ....

ⓐ $_{9}F : 1s^2 , 2s^2 , 2p^5$

ⓑ $_{7}N : 1s^2 , 2s^2 , 2p^3$

ⓒ $_{2}He : 1s^2$

ⓓ $_{3}Li : 1s^2 , 2p^1$

40-

Which of the following represents the electronic configuration of gallium atom $_{31}Ga$ in its excited state ? ....

ⓐ 2, 8, 17, 3

ⓑ 2, 8, 17, 4

ⓒ 2, 8, 18, 3

ⓓ 2, 8, 18, 4

Chapter two

Classification of elements and the long periodic table

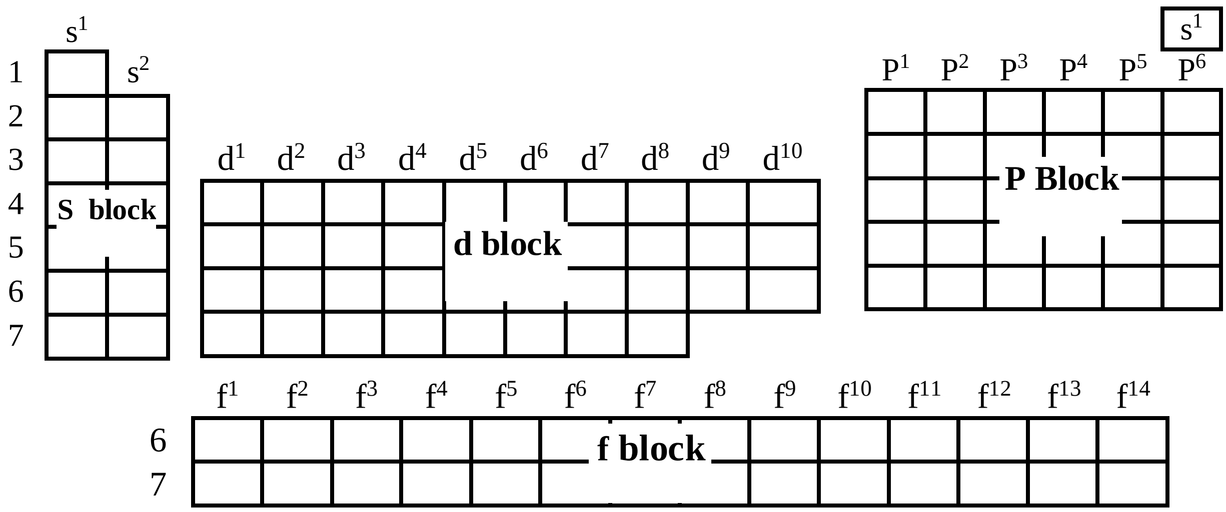
The long form periodic table:
It depends on the building up principle (Auf - bau) The elements are arranged in an ascending order according to their atomic number (each element have one electron more than the element before it) .

The long form periodic table : the elements were arranged in an ascending order according to their atomic numbers

The periodic table is divided into four main blocks

1- S – block elements
2- P – block elements
3- d – block elements
4- f – block elements

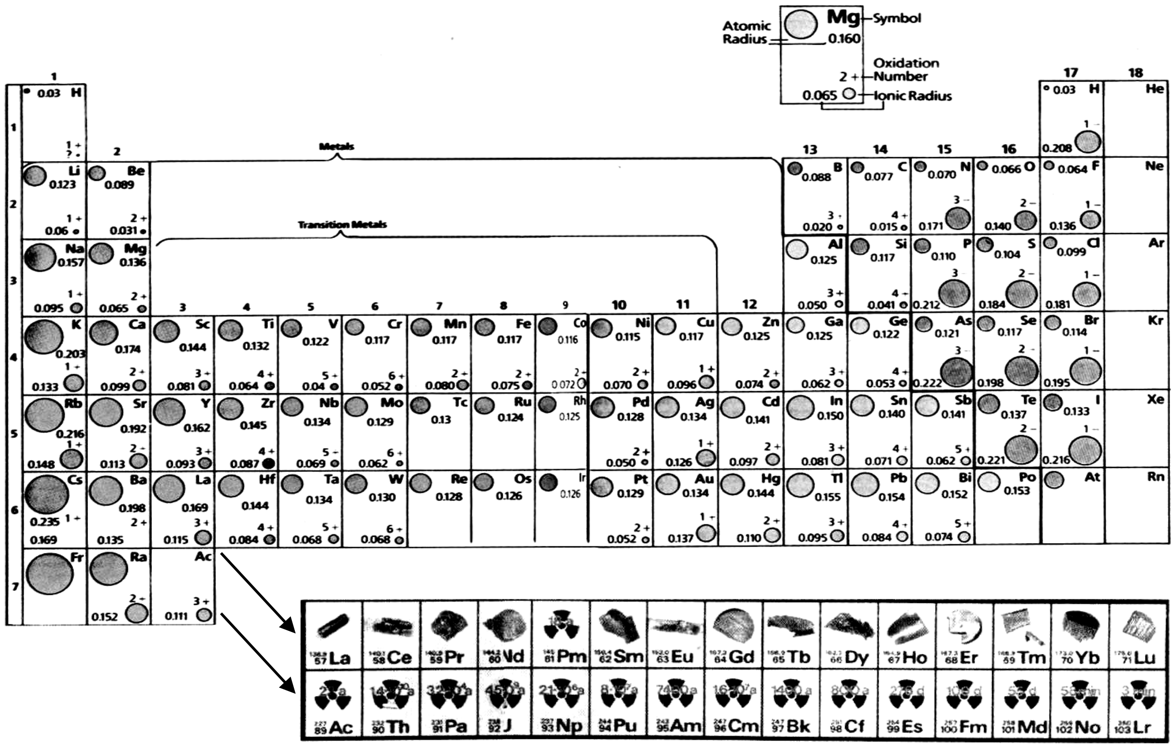
1s	**2s < 2p**	**3s < 3p**	**4s < 3d < 4p**
5s < 4d < 5p	**6s < 4f < 5d < 6p**	**7s < 5f < 6d**	



Sequence of energy sublevels

1S < 2S < 2P < 3S < 3P < 4S < 3d < 4P < 5S < 4d < 5P < 6S < 4f < 5d < 6P < 7S < 5f < 6d

Description of long periodic table :



1- S – block consists of two groups because the S – sublevel consists of one orbital which is filled with two electrons only .

2- P – block consists of (6) groups because the P – sublevel consists of three orbitals which filled with six electrons .

3- d – block consists of (10) groups because the d – sublevel consists of five orbitals which are filled with ten electrons .

4- f block are separated from the table so that the table is not too wide (long)

5- The first period contains two elements because it consists of elements of the sublevel 1S = 2 electrons .

6- The second period contains eight elements because it consists of the sublevel (2S + 2P) = 8 electrons .

7- The third period contains eight elements because it consists of element of sublevel (3S + 3P) = (2 + 6) = 8 electrons .

8- The fourth period contains eighteen elements because it consists of elements of the sublevels (4S + 3d + 4P) = 2 + 10 + 6 = 18 electrons .

9- The fifth period contains (32) elements because it consists of elements of the sublevels (5S + 4d + 5P) = 2 + 10 + 6 = 18 electrons .

10- The six period contains (32) elements because it consists of elements of the sublevels (6S + 4P + 5d + 6P) = 2 + 14 + 10 + 6 = 32 electrons .

<u>How can you find the location and the type of element in the periodic table ?</u>

1- Write the electronic configuration of element in quantum levels .

2- Number of period = The maximum value of principle energy level (quantum number) .

<u>Example :</u>

Find the number of period and group for each of the following element :

Na : The atomic number = 11

Cl : The aomic number = 17

<u>Noble gases</u> : They are the elements of the last column of the P – block all their energy levels are completely filled with electrons .

<u>The representative elements :</u> They are the elements of main group (S – and P – blocks) all their energy levels are completely filled with electrons except for the external energy level .

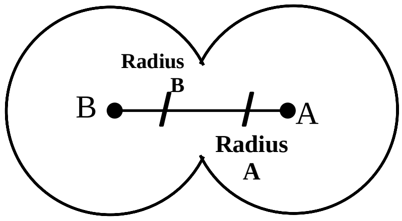
<u>The transition elements</u> : They are the elements of the d – block all their energy levels are completely filled with electrons except for the two external energy levels .

<u>The inner transition elements :</u> They are the elements of the f – block all their energy levels are completely filled with electrons except for the three external levels .

Trends and periodicity of properties in the periodic table

<u>The atomic radius</u> :
We can not determine the atomic radius because electron has a wave motion so it is impossible to determine exactly the location of an electron around the nucleus .

<u>The atomic radius</u> :
It is half the distance between centers of two similar atoms in a diatomic molecule .



<u>The bond length</u> : It is the distance between the nuclei of two bonded atoms .

<u>There are many methods to measure the bond length such as :</u>

1- X – ray . 2- Electron diffraction .

Examples

1- The bond length in the chloride molecule Cl – Cl is 1.98 Å and the length between carbon and chloride atoms C – Cl is 1.76 Å . Calculate the atomic radius of carbon .

<u>solution</u>

The atomic radius of chlorine = $\frac{1.98}{2}$ = 0.99 Å

The atomic radius of carbon = 1.76 – 0.99 = 0.77 Å

3- The bond length in the molecule of NH_3 is 1.0 Å and the bond length in the molecule of H_2 is 0.6 Å . Calculate the bond length in nitrogen molecule (N_2) ?

solution

The atomic radius of hydrogen = $\frac{0.6}{2}$ = 0.3 Å

The atomic radius of nitrogen = 1 – 0.3 = 0.7 Å

The bond length of the N_2 = 0.7 × 2 = 1.4 Å

Atomic radius decrease in period by increasing atomic number ?

Because of increase in the atomic number gradually makes to increase the positive nuclear charge therefore the attractive force of the nucleus for electrons will be increased and atomic radius decreased .

Atomic radius increases in group by increasing of atomic number why ?

Because the increase in atomic number makes to increase the number of energy level screen the attractive force of the nucleus for valence electrons and increase the repulsion force between electrons therefore the atomic radius increases .

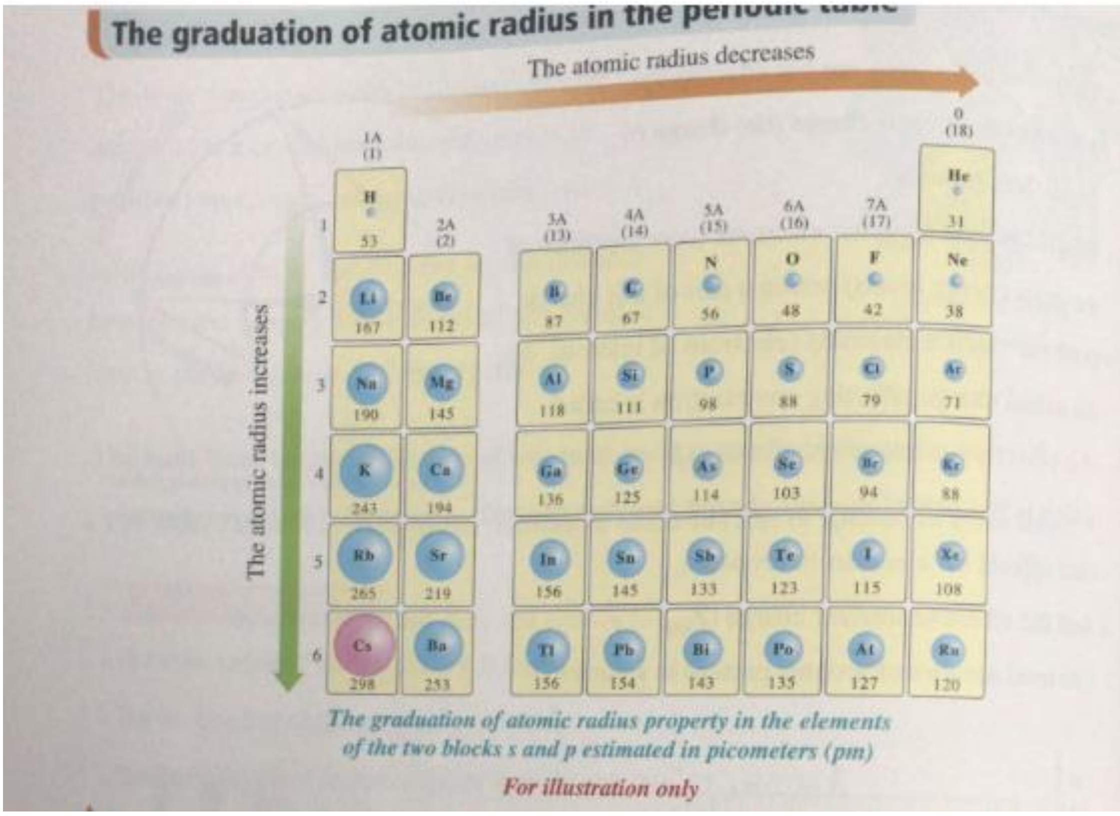
N.B.

1- The cation (+ve ion) radius is smaller than that of its atom ?

This is due to the increasing of positive charges of the protons which attract the valency electrons leading to a decrease in the cation radius .

2- The anion's radius is bigger than that of its atom ?

This is due to the increase in the number of the negative charges in the shells therefore the repulsive force between the electrons increases so the shells move a part and this leads to increasing of the anionic radius than that of its atom .


The graduation of atomic radius in the
The atomic radius decreases
The atomic radius increases
1A (1)
2A (2)
3A (13)
4A (14)
5A (15)
6A (16)
7A (17)
0 (18)
1
2
3
4
5
6
H 53
He 31
Li 167
Be 112
B 87
C 67
N 56
O 48
F 42
Ne 38
Na 190
Mg 145
Al 118
Si 111
P 98
S 88
Cl 79
Ar 71
K 243
Ca 194
Ga 136
Ge 125
As 114
Se 103
Br 94
Kr 88
Rb 265
Sr 219
In 156
Sn 145
Sb 133
Te 123
I 115
Xe 108
Cs 298
Ba 253
Tl 156
Pb 154
Bi 143
Po 135
At 127
Rn 120
The graduation of atomic radius property in the elements
of the two blocks s and p estimated in picometers (pm)
For illustration only

Ionization potential

It is the amount of energy required to remove the smallest bounded electron completely from an isolated gaseous atom

$M \longrightarrow M^{+} + e^{-}$ $\Delta H = +$ ve KJ/Mole

The atom losses electrons and converted into positive ions . It has a positive value (ΔH = + ve KJ/Mole) .

1- **The first ionization energy** : **It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge .**

$M \longrightarrow M^{+} + e^{-}$ $\Delta H = +$ ve KJ/Mole

2- The first ionization energy of noble gas is very high ?

Due to the stability of their electronic configuration because it is difficult to remove an electron from completely filled shell .

3- The ionization energy of element of group (5A)

(N_{7} – P_{15} – As_{33} – Sb_{51} – Bi_{83}) is much greater than any element have the same period because the outer most energy sublevel (P) has three electrons and it is half filled with electrons (nP) and this gives the atom of the element some extra stability so the ionization energy is greater .

4- The ionization energy of sodium is much smaller than that of chlorine ?

Because the atomic size of chlorine is smaller than that of sodium so the attractive force of the nucleus on the valence electrons in the case of chlorine is more strongly and the electrons valence need a higher energy to be separated from the atom .

5- Ionization energy increases period ?

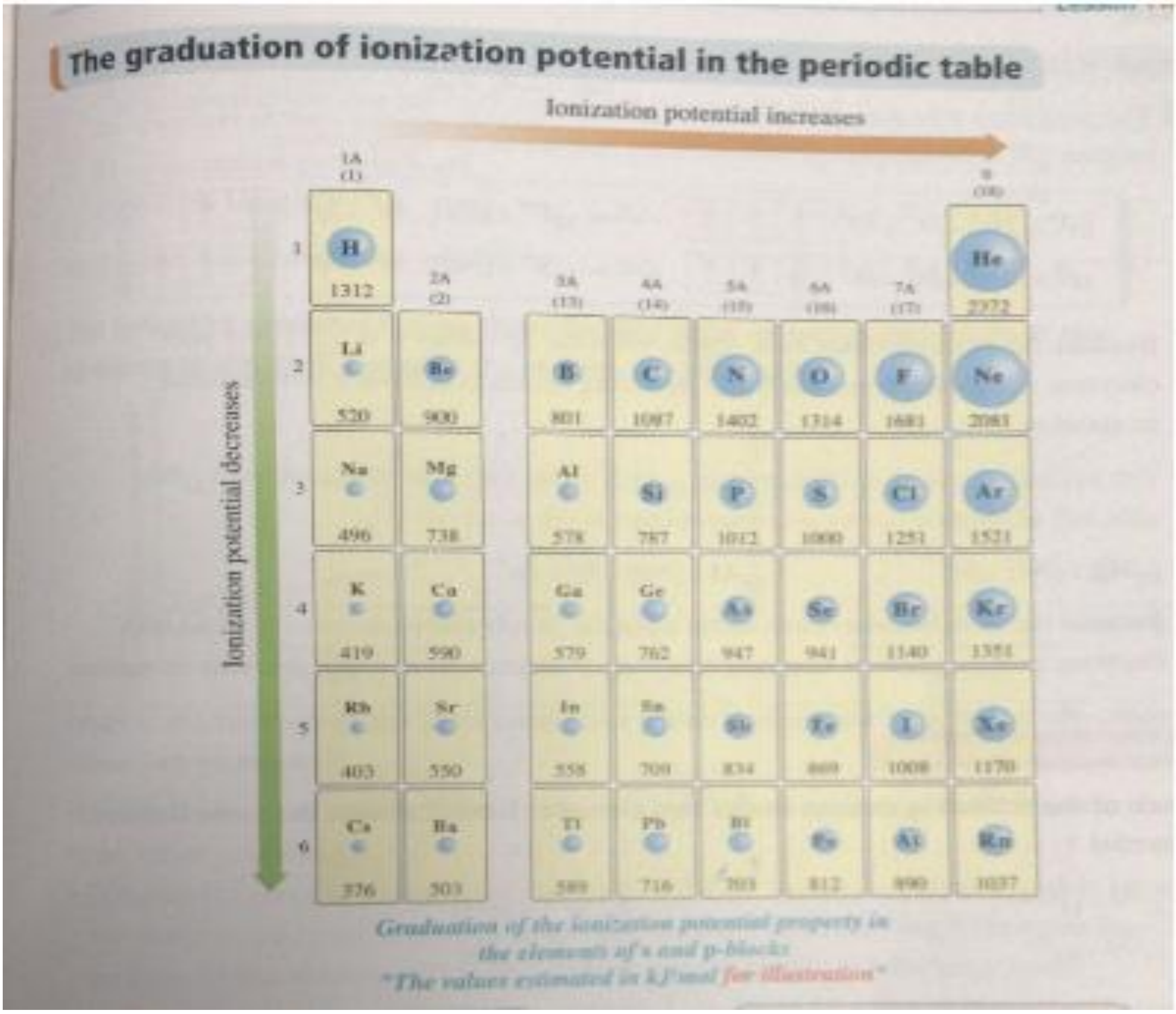
Because the positive nuclear charge gradually increases with the increase of the atomic number led to decrease the atomic radius and increase the attractive force of the nucleus on the valence electrons therefore the electrons needed m large high energy to remove (separated) from the atom .

6- Ionization energy decreases in group ?

Due to the increase in the atomic size and screen of the attraction force of the nucleus on the valence electrons therefore the electrons needed a smaller value of energy to separated from the atom .

7- The ionization energy of elements of group (2A) is much greater than any element have the same period ?

Because the outer most energy level (S) of the element of group (2A) is completely filled with electron (nS^2) and this gives the atom of the element some extra stability so the ionization energy is much greater .



Electron affinity

It is the amount of energy released when an extra electron is added to a neutral gaseous atom to form an anion (- ve ion)

$X + e^- \longrightarrow X^- + \text{energy} \qquad \Delta H = -\text{ ve KJ/Mole}$

G.R.F. In the horizontal periods electron affinity increases with the increase in atomic number ?

Due to the atomic radius (size) gradual decrease so it becomes easier for the nucleus to attract the new electron .

G.R.F. The electron affinity decreases in group ?

Due to the increase of the atomic volume with increase atomic number and this leads to the screeing of the attraction force of the nucleus on the valence electrons .

Exception cases :

Beryllium has a relatively high of electrons affinity due to the stability of its atom that has completely filled orbitals ($1S^2$, $2S^2$) ?

Because the outer most energy sublevel (nS) is completely filled with electron and it gives the atom some extra stability .

Elements of the fifth group (N_7 , P_{15}) have a lower value of electron affinity

Because the outer most energy sublevel (nP) has three electrons and it is half filled with electrons it gives the atom some extra stability (N_7 : $1S^2$, $2S^2$, $2P^2$).

Noble gases have not (small) electron affinity

Because all energy sublevels are completely filled with electrons which gives the atoms great stability .

Electron affinity of Fluorine (F_9) is less than that of chlorine (Cl_{17}) ?

Because the atomic radius (size) of fluorine atom is smaller than that of chlorine atom and when fluorine atom gains electron it is affected by a great repulsion force bigger than that in chlorine atom and fluorine atom is very small size .

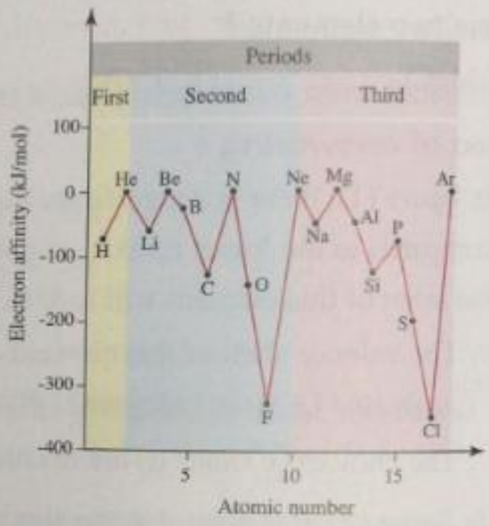
The graduation of electron affinity in the periodic table

Electron affinity increases →

Electron affinity decreases ↓

	1A	2A	3A	4A	5A	6A	7A	0
1	H -73							He >0
2	Li -60	Be >0	B -27	C -122	N >0	O -141	F -328	Ne >0
3	Na -53	Mg >0	Al -43	Si -134	P -72	S -200	Cl -349	Ar >0

The values of the electron affinities of the first 18 elements in the periodic table "in kJ/mol"



In the same period

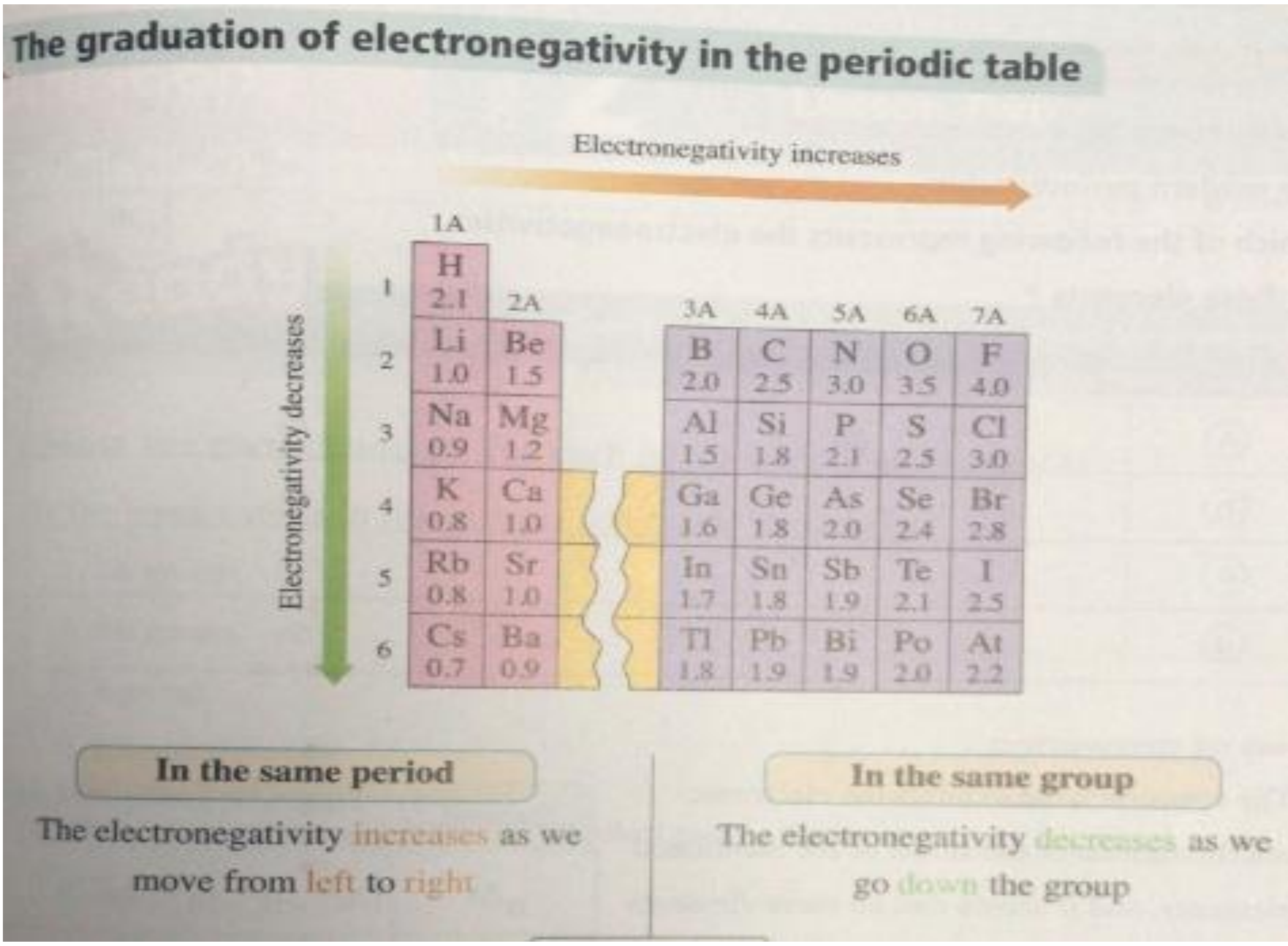
The electron affinity increases as we move from left to right

In the same group

The electron affinity decreases as we go down the group

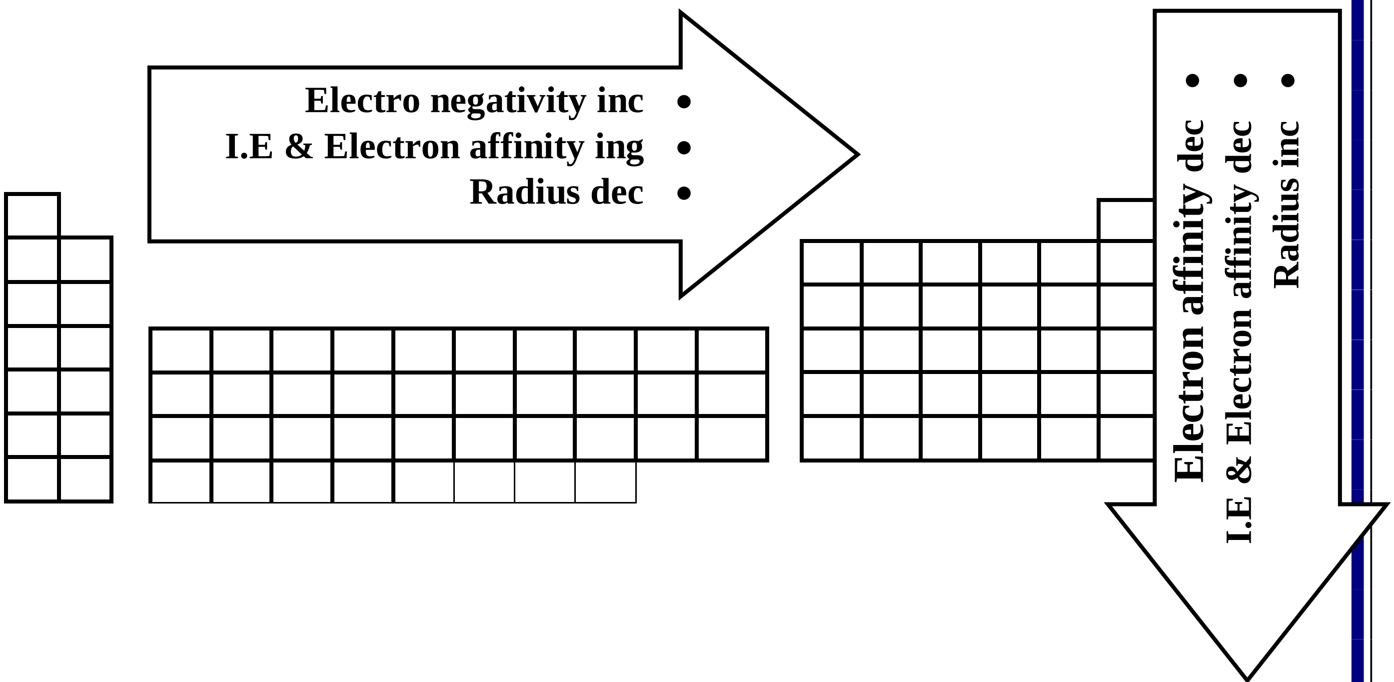
Electro negativity

The tendency of an atom to attract the electrons of chemical bond to itself .
It is the querage of the ionization potential and electron affinity .



Compare between ionization energy , electron affinity and electro negativity

Ionization energy	Electron affinity	Electro negativity
It is the amount of energy needed to remove the least connected electron bond in a single atom	It is the amount of energy released when an extra electron is added to a neutral single atom to form an ion	It is tendency of an atom to attract the electrons of chemical bond to itself
It refers to the atom in its single state	It refers to the atom in its single state	It refers to the atoms which linked together in the molecule
It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius	It is inversely proportional to the atomic radius
The atom losses electrons and converted into positive ion	The atom gains electrons and convers into negative ion	
It has a positive value	It has a negative value ΔH = - value type of reaction is exothermic	



Differentiation between metal and non metal

Metals	Non metals
Valence shell has less than help its capacity of electrons (1 or 2 or 3)	Valence shell has more than half its capacity of electrons (5 or 6 or 7)
They one called electropositive elements .	They are called electronegative elements
They have Relatively large atomic radius therefore the ionization energy and electron affinity and electro negativity have smell values	They have small atomic radius ionization energy and electro negativity and electron affinity have high value
They have a good conductivity of electricity .	they have a bad conductivity of electricity .

N.B.

1- Fluorine is the strongest non metal while caesium is the strongest metals Because fluorine has smallest radius while caesium has the biggest radius .

2- Metals are considered as electropositive elements because metals lose electrons to form positive ions .

3- Non metals are considered as electronegative elements because non metals gains electrons to from negative ions .

4- Metal are good conductors of electricity because they have few valence electrons which can transfer easily from one position to another in the metal structure .

5- Non metals are bad conductors of electricity because their valence electrons are strongly bounded to the nucleus due to the small atomic size therefore it is difficult for the valency electrons to be transferred .

6- Metals have small values for ionization energy and electron affinity because they have large atomic radius .

7- The strongest metal lie at the bottom on the left and side of the periodic table because of the increasing in the atomic number the atomic radius increases gradually so the attractive force of the nucleus to the valency electrons decreases therefore it is very easy for the atom to lose the valency electrons so the metalic property increases .

Metalloids

1- They are elements whose valency shell contains 4 electrons .
2- Sometimes they act as metal (when they gains electrons) .
3- They act as semiconductors (Boron – silicon) which are used in or transistors and knows as semiconductors electronic instruments .

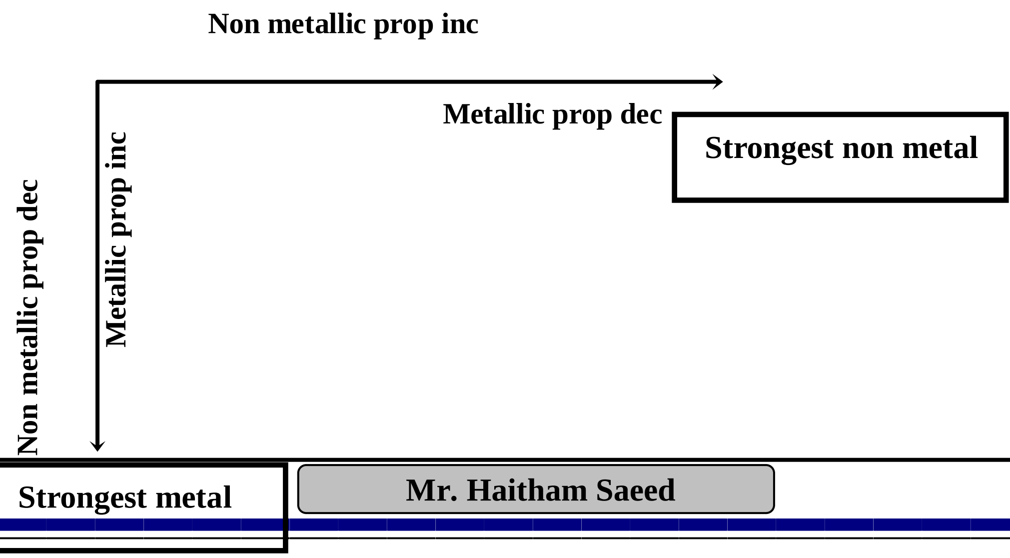
metals $\xrightarrow[\text{in period}]{\text{decreases}}$ metalloids $\xrightarrow[\text{in period}]{\text{increase}}$ non metal

* Because the atomic size decrease and the attractive force of nucleus on the valence electron will be increased therefore it is difficult for the atom to lose the valence electron .

<u>Metalloids increase in period</u> because the atomic size decrease and the attractive force of nucleus to electron will be increased therefore it becomes easier for the nucleus to gain a new electron .

<u>In groups</u> : Metals increase in group because the increase in the atomic number makes to increase of energy levels and screen (decrease) the attractive force of the nucleus on the valence electrons therefore it is easy for the atom to lose the valence electrons .

In groups : non metals decrease because the increase in the atomic number makes to increase the energy level and screen the attractive force of nucleus on the valence electrons therefore it is difficult for the atom to gain a new electron .
Fluorine is considered as the strongest non – metal because the atomic size of fluorine is very small therefore the attractive force of the nucleus to the electrons will be increased therefore it is very easy for the atom to gain a new electrons .



Acidic – basic properties

Acidic oxides	Basic oxides
They are non – metalic oxides such as CO_2 , SO_2 , SO_3 , P_2O_5	They are metalic oxides such as Na_2O , K_2O , MnO , CaO , BaO
They dissolve in water to form acids $Co_2 + H_2O \longrightarrow H_2CO_3$ $SO_2 + H_2O \longrightarrow H_2SO_3$ $SO_3 + H_2O \longrightarrow H_2SO_4$ $P_2O_5 + H_2O \longrightarrow 2H_3PO_4$	Some basic oxides dissolve in water to form alkalies and others are not $Na_2O + H_2O \longrightarrow 2NaOH$ $K_2O + H_2O \longrightarrow 2KOH$ $CaO + H_2O \longrightarrow Ca(OH)_2$ $MgO + H_2O \longrightarrow Mg(OH)_2$
They react with alkalis to form salt and water $CO_2 + 2NaOH \longrightarrow Na_2CO_3 + H_2O$ $SO_2 + 2NaOH \longrightarrow Na_2SO_4 + H_2O$	They react with acids to form salt and water $Na_2O + 2Hcl \longrightarrow H_2O + 2Nacl$ $MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
They do not reaction with acids	They do not react with alkalis

The oxyacids : are acids that contain hydrogen, oxygen and a third element usually a non-metal .

- It can take the following symbols: - M On (OH)m

where , (M) is the atom of the element .
(n) is the number of oxygen atoms .
(m) is the number of hydroxyl groups .

The strength of oxyacids are depends on the number of oxygen atoms which does not linked with hydrogen atoms when this number increase , the strength of the acid .

Acid Name	The number of free atoms of Oxygen	Acidic Property
H_4SiO_4	-	Weak acid
H_3PO_4	1	Moderate acid
H_2SO_4	2	Strong acid
$HClO_4$	3	Very strong acid

Amphoteric oxides :

1- They react with bases and acids to form (give) salt and water in both cases .

2- They have both acidic and basic properties such as Al_2O_3 , ZnO , Sb_2O_3 , SnO .

How can you obtain CuO from mixture of (CuO + Al_2O_3) ?

3- KOH is strong base than of NaOH because the atomic radius of K is greater than that of Na so the bond strength between K and hydroxide group (OH^-) decreases therefore it is easy to separate hydroxide ions to form strong base .

4- HF is weak acid because the atomic size of fluorine atom is very small so it has ability to attract the electrons bond to itself therefore it is difficult to separate hydrogen ions to form strong acid .

5- HI is stronger acid because the atomic (size) radius of iodine atom is largest to it has a weak ability to attract the electron bond to itself therefore it is easy to separate hydrogen (H) ions to form strong acid .

Oxidation number : at last we defined it as it is the number of hydrogen atoms that combine with or can be replaced by an atom of the element .

The modern definition : the number of single (unpaired) electrons in the valence shell of the atom .

N_7	**$1S^2$**	**$2S^2$**	**$2p^3$**	**Trivalet**
O_8	**$1S^2$**	**$2S^2$**	**$2p^4$**	**Divalent**
F_9	**$1S^2$**	**$2S^2$**	**$2p^5$**	**Mono valet**

Oxidation number : it is a number that refers to the electric charge (+ve or -ve) that atom would have in the compound .

Rules for assigning oxidation numbers

Oxidation number of Oxygen = 2- in most of its compounds except in peroxides e.g. (hydrogen peroxide H_2O_2) , sodium peroxide Na_2O_2 and potassium peroxide K_2O_2 is (1-) supper oxide e.g. KO_2 the oxidation number of oxygen is (½-)

The oxidation number of oxygen = 2+ in OF_2 because the electro negativity of fluorine is higher than the electron negativity of oxygen .

The oxidation number of hydrogen in its hydrides = 1- such as NaH , KH , CaH_2 .

The oxidation number of any chemical compound = zero because the algebraic sum of the oxidation number of its atom = zero .
Such as NaCl , HCl , CuSO4

The oxidation number of any element in its pure state = zero .
e.g. Fe , Cl , Na , O_2S , P4 .

The oxidation number for any atomic group (poly atomic ion) = the number of charge on the group (ion) such as OH^- , $SO4^{--}$, $CO3^{--}$, $NH4^+$, $PO4^{---}$.

Examples

Calculate the oxidation number of chlorine in these compound , and for sulphure

Nacl **Nacl = zero** **+1 +Cl = 0** **Cl = -1**	**$NaClO_2$** **$NaClO_2$ = zero** **1+Cl 2-×2 = zero** **Cl = +3**	**$NaClO_4$** **$NaClO_4$= zero** **+1 Cl -8 = 0** **Cl = +7**	**H_2S = zero** **2×1+S= zero** **S = 2-**	**$SO3^{--}$** **SO3 = -2** **S -6 = -2** **S = 4**
K_2S **K_2S = zero** **C + S = 0** **S = -2**	**$Na_2S_2O_3$** **$Na_2S_2O_3$ = zero** **2×1+S_2 -6= zero** **S = 2**	**S = zero**	**H_2SO_4= zero** **2 + S-8= 0** **S = 6**	**$SO4^{--}$** **SO4 = -2** **S -8 = -2** **S = +6**

<u>The oxidation process</u> : it is the process of losing electrons due to increase the oxidation number for the element .

<u>The Reduction process</u> : it is the process of gaining electrons due to decrease the oxidation number for the element .

Oxidation	Reduction
Zero ⟶ (+) value	(+) value ⟶ zero
(+) small value ⟶ (+) big value	(+) big value ⟶ (+) small value
(-) value ⟶ zero	zero ⟶ (-) value
(-) big value ⟶ (-) small value	(-) small value ⟶ (-) big value

Explain the type of change (oxidation or Reduction for carbon in this Reaction .

$C_2H_5OH \longrightarrow CH_3CHO \longrightarrow CH_3COOH \longrightarrow CO_2$

$C = -2 \xrightarrow[\text{Oxidation}]{} C = +1 \longrightarrow C = \text{zero} \longrightarrow C = +4$

Oxidation Oxidation Oxidation

Explain the type of change (oxidation or reduction) that of chromium and iron in this reaction .

$K_2CR_2O_7 + 6FeCl_2 + 14HCl \longrightarrow 2KCl + CrCl_3 + 6FeCl_3 + 7\ H_2O$

Solution

$K_2Cr_2O_7 \longrightarrow CrCl_3$ **$FeCl_2 \longrightarrow FeCl_3$**

$Cr = +6 \xrightarrow{\text{Reduction}} Cr = +3$ **$Fe = +2 \xrightarrow{\text{Reduction}} Fe = +3$**

R

Gradation of oxidation number in periodic table :

Group	I	II	III	I V	V	V I	V II
Element	Na	Mg	Al	Si	P	S	Cl
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Oxidation Number	1 +	+ 2	+ 3	+ 4	+ 5	+ 6	+ 7

In the case of hydride

- The oxidation number of element (metal) of 1st and 3rd groups in their compounds agrees with group number (it takes positive oxidation number to which they belong)
- Most elements in the middle of the table have variable oxidation number in different compounds .

The oxidation number of the elements in groups from IV to VII = the group number -8 = - value because the elements gain electrons to complete a full shell .

Group	**I**	**II**	**III**	**I V**	**V**	**V I**	**V II**
Element	Li	Be	B	C	N	O	F
Hydride	LiH	BeH_2	BH_3	CH_4	NH_3	H_2O	HF
Oxidation Number	1 +	+ 2	+ 3	+ 4	- 3	- 2	- 1

The oxidation number of noble gases of (group zero) is zero because they do not combine to form compounds .

Questions on Chapter 2

A- Choose the best answer:

1- The sublevel (d) contains .... orbitals, it can accommodate up to .... electron . (2 - 3 - 5 - 10 - 14)

2- An element of atomic number 10, the number of sublevels filled with electrons are .... (2 - 3 - 5)

3- The valency shell of the atoms of fifth group elements contain ....

a) np4 b) np5 c) np d) np3

4-

The properties of the two elements .... are similar.

(a) $_{19}K$, $_{20}Ca$

(b) $_{31}Ga$, $_{32}Ge$

(c) $_{17}Cl$, $_{35}Br$

(d) $_{55}Cs$, $_{56}Ba$

5- Group 5-A elements are classified as .... block . (S-P-d-f)

6-

What is the number of the periods in the periodic table in which the elements from hydrogen ($_1H$) to argon ($_{18}Ar$) are located ?

(a) 2 periods.

(b) 3 periods.

(c) 4 periods.

(d) 8 periods.

7-

The elements which follow neon gas ($_{10}Ne$) and precede rubidium element ($_{37}Rb$) are located in ....

(a) the third period only.

(b) the fourth period only.

(c) the third and the fourth periods.

(d) the fourth and the fifth periods.

8-

What is the type of the elements whose last electronic configuration is : $ns^{1:2}$, $np^{1:5}$?

(a) Representative.

(b) Main transition.

(c) Inner transition.

(d) Noble.

9-

Which of the following choices represents the electronic configuration of an alkaline earth metal ?

(a) [Ar] , $4s^1$, $3d^5$

(b) [Ar] , $4s^2$, $3d^6$

(c) [Rn] , $7s^2$

(d) [Xe] , $6s^2$, $5d^1$, $4f^7$

10-

All the following are among the properties of the elements $_4Be$, $_{12}Mg$ and $_{20}Ca$, except that ....

(a) the last sublevel s contains 2 electrons.

(b) the sublevel p in the valence shell contains a pair of electrons.

(c) they are representative elements.

(d) they are located in group (2A).

11-

What is the type of the element which contains 2 electrons in the sublevel whose value of the quantum number (ℓ) is 2 ?

(a) Main transition.

(b) Inner transition.

(c) Noble.

(d) Representative.

12-

An element with atomic number 42, the number of its half filled orbitals is ....

(a) 1 (b) 4

(c) 5 (d) 6

13-

The electronic configuration of ruthenium ion $_{44}Ru^{3+}$ is ....

ⓐ [Kr] , $4d^3$, $5s^2$

ⓑ [Kr] , $4d^6$, $5s^2$

ⓒ [Kr] , $4d^5$

ⓓ [Kr] , $4d^6$

14-

What is the atomic number of the element (X) whose electronic configuration ends with : $ns^1, (n-1)d^5$ and its electrons are distributed in 5 principal energy levels ?

ⓐ 29

ⓑ 24

ⓒ 47

ⓓ 42

15-

If the electronic configuration of an element atom is [Xe], $6s^2, 4d^1, 5f^7$

Which of the following choices represents the distribution of the electrons in the principal energy levels ?

ⓐ 2 – 8 – 18 – 32 – 4

ⓑ 2 – 8 – 18 – 18 – 8 – 2

ⓒ 2 – 8 – 18 – 25 – 9 – 2

ⓓ 2 – 8 – 18 – 32 – 4

16-

Magnesium ion $^{24}_{12}Mg^{2+}$ contains ....

ⓐ 12 protons, 10 electrons.

ⓑ 24 protons, 26 electrons.

ⓒ 12 protons, 13 electrons.

ⓓ 24 protons, 14 electrons.

17-

The highest number of the unpaired electrons is in .... [Atomic number of iron is : 26]

(a) Fe
(b) Fe^{4+}
(c) Fe^{2+}
(d) Fe^{3+}

18-

Which of the following relations is correct for the elements of the same period ?

(a) The radius of M^+ ion > That of X^- ion.
(b) The radius of X^- ion > That of X atom.
(c) The radius of M^+ ion = That of X^- ion.
(d) The radius of M^+ ion > That of M atom.

19-

If the atomic radius of rubidium is 253 Pm

What is its ionic radius (rounded to the nearest integer) ?

(a) 148 Pm
(b) 253 Pm
(c) 275 Pm
(d) 300 Pm

20-

Which of the following elements has the lowest second ionization potential ?

(a) $_{16}S$
(b) $_{11}Na$
(c) $_{7}N$
(d) $_{5}B$

21-

Which of the following has a higher value in lithium Li than in potassium K ?

(a) First ionization potential.
(b) Atomic radius.
(c) Atomic number.
(d) Ionic radius.

22-

Nitrogen gas is less active than fluorine gas, because ....

(a) the boiling point of nitrogen is less than that of fluorine.

(b) the molar mass of nitrogen is less than that of fluorine.

(c) the atomic radius of nitrogen is larger than that of fluorine.

(d) the electronegativity of nitrogen is higher than that of fluorine.

23-

What is the substance which dissolves in water forming an alkaline solution ?

(a) MgO

(b) Al_2O_3

(c) SiO_2

(d) SO_2

24-

The opposite table shows some elements of the third and fourth periods in the periodic table.
What is (are) the element(s) of the fourth period whose oxide(s) dissolve(s) in water forming acidic solution ?

3rd period	Al	Si	P	S
4th period	Ga	Ge	As	Se

(a) As and Ga

(b) Ga and Ge

(c) Ga and Se

(d) Se only.

25-

What is the formula of the oxide of the element (M) which is located in group (3A) in the periodic table ?

(a) M_2O_3

(b) M_3O_2

(c) MO

(d) M_3O_4

26-

The weakest oxygenated acid in the fourth period in the periodic table is ....

(a) $Ge(OH)_4$

(b) $BrO_3(OH)$

(c) $AsO(OH)_3$

(d) $SeO_2(OH)_2$

27-

Perchloric acid is a ....

(a) monohydric acid.

(b) dihydric acid.

(c) trihydric acid.

(d) tetrahydric acid.

28-

Element (M) is located in group (5A).

What is the probable hydroxy formula of its oxygenated acid ?

(a) $M(OH)_4$

(b) $MO(OH)_3$

(c) $MO_2(OH)_2$

(d) $MO_3(OH)$

29-

From these equations :

- $Pb(OH)_{2(s)} + 2HNO_{3(aq)} \longrightarrow Pb(NO_3)_{2(aq)} + 2H_2O_{(l)}$
- $Pb(OH)_{2(s)} + 2NaOH_{(aq)} \longrightarrow Na_2PbO_{2(aq)} + 2H_2O_{(l)}$

It is concluded that ....

(a) the basic property of $Pb(OH)_2$ is stronger than its acidic property.

(b) the aqueous solution of $Pb(OH)_2$ is amphoteric.

(c) $Pb(OH)_2$ as an acid is stronger than its strength as a base.

(d) (Pb – O) bond strength is equal that of (O – H) bond.

30-

In which of these compounds does nitrogen have two oxidation numbers ?

(a) $NaNO_3$

(b) NH_4NO_3

(c) NH_4Cl

(d) NH_2NH_2

31-

What is the oxidation number of phosphorus in pyrophosphate ion $(P_2O_7)^{4-}$?

(a) +3.5
(b) +5
(c) +7
(d) +10

32-

When $(MnO_4)^-$ reacts and is converted to Mn^{2+} , $(MnO_4)^-$ is ....

(a) reduced, as the oxidation number of manganese increases.
(b) oxidized, as the oxidation number of manganese increases.
(c) reduced, as the oxidation number of manganese decreases.
(d) oxidized, as the oxidation number of manganese decreases.

33-

In which of the following changes an oxidation process occurs to vanadium V ?

(a) $VO_2 \longrightarrow V_2O_3$

(b) $V_2O_5 \longrightarrow VO_2$

(c) $V_2O_3 \longrightarrow VO$

(d) $V_2O_3 \longrightarrow V_2O_5$

34-

In the following oxidation-reduction reaction :

$$2Cr^{3+}_{(aq)} + 3Cl_{2(aq)} + 7H_2O_{(l)} \longrightarrow Cr_2O^{2-}_{7(aq)} + 6Cl^-_{(aq)} + 14H^+_{(aq)}$$

Which of the following loses electrons ?

(a) Cl_2

(b) Cr^{3+}

(c) H_2O

(d) $Cr_2O_7^{2-}$

35-

In the opposite redox reaction : $Fe^{3+} + Al \longrightarrow Fe + Al^{3+}$

The electrons transfer from ....

(a) $Fe^{3+} \longrightarrow Al$

(b) $Al \longrightarrow Fe^{3+}$

(c) $Fe \longrightarrow Fe^{3+}$

(d) $Al^{3+} \longrightarrow Al$

36-

In the second period, on moving from lithium to fluorine, ...

(a) the atomic size decreases.

(b) the ionization potential decreases.

(c) the electronegativity decreases.

(d) the nuclear charge decreases.

37-

What are the two metals whose oxides can react with both acids and alkalis ?

(a) Na , Zn

(b) Mg , Al

(c) Mg , Be

(d) Al , Zn

38-

The electronegativity of aluminum $_{13}Al$ is similar to the electronegativity of ....

ⓐ barium $_{56}Ba$

ⓑ beryllium $_{4}Be$

ⓒ magnesium $_{12}Mg$

ⓓ strontium $_{38}Sr$

39- The oxidation number of hydrogen in hydrogen molecule is (1, 2, 3, 0) and its oxidation number of hydrogen chloride is (1, 2, 3, 0).

40- The oxidation number of oxygen in H_2O_2 is ....

a)+2 b)+l c)-2 d)-l